Unit B: Atoms and Elements (Physical Science: Chemistry)

Chapter 4 - People of all cultures investigate matter according to various properties

Outcomes

- Distinguish between physical and chemical properties of common substances, including those found in household, commercial, industrial, and agricultural applications

4.1 Investigating Matter

- Three models for understanding matter include:
  - Indigenous
    - Medicine Wheel (wind / fire / rock / water)
  - Ancient Greek
    - elements (air / fire / earth / water)
  - Scientific

What is Matter? Matter is anything that has mass and takes up space.

- Composed of particles. The type and arrangement of the particles determine its properties
- Matter is composed of particles:

Pure Substance – contains only one kind of particle
Mixture – contains at least two different pure substances
Suspension – particles of one substance are held within another
Solutions – a mixture where the different parts are not visible
Heterogeneous Substance – a mixture where each part can be seen (mechanical mixture)
Element – pure substance that can not be broken down into simpler substances
Compound – pure substance that contains two or more elements in fixed proportion
Atom – smallest particle of any element
Molecule – contains two or more atoms

Read p. 121 – 127
Learning Checkpoint # 1 – 3 (p. 124)
4.1 Check and Reflect # 1 – 9 (p. 127)
Safety Symbols / B1 & B2

4.2 Physical and Chemical Properties

Physical Properties

A physical property is a characteristic or description of a substance that may help to identify it by observation or measurement.
Properties identified by the senses:

Colour	Texture
Odour	Lustre
Clarity	Taste

States of Matter / Phases of Matter

Solid
Shape – definite, has a fixed shape
Volume – definite, has a fixed volume
Liquid
Shape – indefinite, takes the shape of its container
Volume - definite, has a fixed volume
Gas
Shape - indefinite, takes the shape of its container
Volume – indefinite, fills the entire container

Other Physical Properties
Hardness – a measure of a solids resistance to being scratched or dented
Example: diamond is hard, chalk is soft
Malleability – a measure of a solids ability to be hammered or bent into different shapes
Example: aluminium is malleable, glass is brittle
Ductility - a measure of a solids ability to be pulled into wires
Example: Cu/Al is ductile, coal is not
Melting and Boiling Points – the temperature at which substances change states
Example: m.p. of water is 0°C;
b.p. of water is 100°C
Crystal Form – many minerals form characteristic structures
Example: salt is cubic, plastic has no crystal form
Solubility – the ability of a solid to dissolve in a solvent (water)
Example: salt is soluble in water, iron is not soluble
Viscosity – how a liquid flows, the thicker the more viscous
Example: maple syrup is more viscous than water
Density – the amount of matter per unit volume
Example: lead is denser than feathers; density of water is 1.00 g/cm^3

Chemical Properties

A chemical property describes the chemical reactivity and the behaviour of a substance as it becomes a new substance.
Examples

- Absorbs heat
- Forms gas when heated

- Emits heat energy during reaction
- Emits light during the reaction
- Forms a precipitate (solid) in a solution

Also;
Combustibility – If a substance is combustible or flammable it will burn when exposed to a flame.
Example: gasoline is flammable, water is not
Reaction with Acid/Water – Some substances react with acids while some do not.
Example: Mg reacts with acid, gold does not

Chemical and Physical Changes

Physical Changes
In a physical change, the substance involved remains the same substance, even though it may change state or form.
Chemical Changes
In a chemical change, the original substance is changed into one or more different substances that have different properties.

Read p. 131 – 141
Learning Checkpoint # 1 – 5 (p. 137)
4.2 Check and Reflect # 1 – 7 (p. 141)
Chapter 4 Review # 1 – 23 (p. 146 – 147)

Chemical Symbols and Formula

Chemical Symbol – an abbreviation of the name of an element
Examples:

- Carbon - C
- Tungsten – W
- Iron - Fe

Chemical Formulas – combination of symbols that represent a particular compound
Examples:

- hydrogen peroxide – H_2O_2
- water – H_2O
- acetic acid (vinegar) – $C_2H_4O_2$

Chapter 5 - The periodic table is a tool for organizing scientific understanding of elements

Outcomes

- Analyze historical explanations of the structure of matter up to and including the models of Dalton, Thomson, Rutherford, and Bohr
- Demonstrate an understanding of the classification of pure substances (elements and compounds), including the development and nature of the Period Table

5.1 Developing Models of Matter

The Particle Theory

1. All matter is made up of tiny particles.
2. All the particles of one substance are the same. Different substances are made of different particles.
3. Particles are always moving the more energy they have the faster they move.
4. There are attraction forces between the particles. The closer they are the stronger the force.

Models For Atoms

Developing Models of Matter (Time Line)

450 B.C.E Greeks believed matter was composed of four elements Earth/Air/Fire/Water

400 B.C.E Democritus (Greek) suggested that all matter is composed of tiny particles called atoms (indivisible)

350 B.C.E Aristotle (Greek) believed the four element model and because his influence was so great it was the accepted theory

500 – 1600 A.D Alchemists (Pb into Au) devised many element symbols but still believed the four element theory

1650

Boyle (English) did not believe the four element theory and defined an element as a pure substance that can not be broken down into simpler substances

1700's

Priestly isolates oxygen.

Lavoisier concludes air is a mixture containing oxygen

Cavendish mixes an acid with metal and isolates hydrogen

Cavendish than mixes hydrogen and oxygen and made water disproving the thought that water was an element

1808 – Dalton

1897 – Thomson

1904 – Nagaoka

1914 – Rutherford

1921 – Bohr

- Models have changed based upon experimental evidence obtained over time

An element is a pure substance made up of one type of particle or atom that cannot be broken down into other substances by chemical reactions

Read p. 149 – 155

Learning Checkpoint # 1 – 4 (p. 154)

5.1 Check and Reflect # 1 – 14 (p. 155)

Inside the Atom

Subatomic Particles
Protons (p^+) – positive charge, mass of 1, located in the nucleus
Neutrons (n^0) – no charge, mass of 1, located in the nucleus
Electrons (e^-) – negative charge, mass of 1 / 2000th of a proton, located outside the nucleus in the electron cloud.
Nucleus – area in the center of the atom, smallest and heaviest part of the atom, positive charge
Electron Cloud – area outside of the nucleus where the electrons are located, negative charge

Diagram of the Atom

Counting Subatomic Particles

Introduction - Notation
Standard Atomic Notation
35 – *mass number*

Cl – *chemical symbol*

17 – *atomic number*
Atomic Number – number on the top left of the periodic table
Mass Number – the number of protons and neutrons added together

Periodic Table

Atoms
Using the Periodic Table of Elements
# of Protons - always equal to the elements atomic number
# of Neutrons – equal to the rounded atomic mass (to a whole number) minus the # of protons
# of Electrons – equal to the number of protons for all atoms

Diagram for Atoms

Isotopes

Isotopes – an element can have atoms with differing numbers of neutrons they are called isotopes.
Chlorine – 35 and Chlorine – 37 are isotopes of chlorine
Chlorine – 35 has 18 neutrons while
Chlorine – 37 has 20 neutrons

Counting Particles for Isotopes

# of Protons - always equal to the elements atomic number
# of Neutrons – equal to the mass number minus the # of protons
# of Electrons – equal to the number of protons for all isotopes

Diagram for Isotopes

Ions

Atoms can gain or lose electrons to become charged. Charged atoms are called ions. The number of protons and neutron do NOT change.

mass number – **35** **-1** – *charge*

Cl – *chemical symbol*

17 – *atomic number*

If an atom gains electrons (negatives) it becomes negatively charged and is called an anion.
If an atom loses electrons (negatives) it becomes positively charged and is called a cation.
To calculate the number of electrons in an ion you must add or subtract the charge.

Counting Particles for Ions

# of Protons - always equal to the elements atomic number
# of Neutrons – equal to the rounded atomic mass (to a whole number) minus the # of protons
Electrons in an Anion – # of protons plus the charge number
Electrons in a Cation – # of protons minus the charge number

Diagram of an ion

Counting Particles Assignment

Bohr-Rutherford Models

In Bohr-Rutherford diagrams a circle drawn in the center represents the nucleus and contains the number of protons and neutrons. Electrons are drawn as dots on circular orbits around the nucleus.

The number of electrons that are held in each circle or orbit is defined. The first orbit holds 2, the second orbit hold 8, the third orbit holds 8, and the fourth orbit hold 2. Bohr-Rutherford diagrams are only accurate for the first twenty elements.

5.2 The Elements

- Scientists classify elements as metals, non-metal, or metalloids based on chemical properties
- Each element has a standard name and symbol which consists of one or two letters

Read p. 157 – 165
Learning Checkpoint # 1 – 4 (p. 162)
5.2 Check and Reflect # 1 – 9 (p. 165)

5.3 Combining Elements to Form Compounds

- Compounds are pure substances composed of two or more elements that combine chemically in a specific ratio
- Ionic compounds form when metallic and non-metallic elements form chemically

Ionic Compounds

- Consist of at least one metal and one non-metal
- Created by an atom or group of atoms losing or gaining electrons
- Metals lose electrons
- Non-metals gain electrons

Example

How Elements form Compounds

1. Metals combine with non-metals in many compounds
2. Write the name/symbol of the metal first and the non-metal second
3. When naming change the ending of the non-metal to –ide
4. Each atom has its own combining capacity
5. Atoms combine so that each can fill its combining capacity

Read p. 167 - 170
Learning Checkpoint p. 171 # 1 - 4

Molecular Compounds

Created when two or more non-metals combine
Consist of a covalent bond (a shared pair of electrons)

Read p. 171 – 173
5.3 Check and Reflect # 1 – 12 (p. 173)

5.4 The Periodic Table

- The PT of E organizes the elements by their properties, such as mass and melting point
- Atomic number is the number of protons in the nucleus and is unique for each element
- Bohr Diagrams of the 1st 20 elements reveal important patterns that relate to an elements properties

The Periodic Table was constructed by a Russian scientist named Dmitri Mendeleev. Mendeleev created a periodic law that states:

"If the elements are arranged according to their atomic mass, a pattern can be seen in which similar properties occur regularly"

Element Song

Groups of Elements
The periodic table has columns called groups and rows called periods. The elements are arranged in order of increasing atomic number from left to right in each period.

Groups
IUPAC #1 – 18
Old System Roman Numerals with an A or B
Elements with similar chemical properties
Same number of electrons in the outer most orbit, called valence electrons
Representative Elements Roman numeral = valence electrons
Sometimes called "Chemical Families"

Periods
#1 – 7 Top to Bottom
Each period represents an orbit or electron shell

Classification Systems

Metals, Non-metals, and Metalloids
Representative Elements, Transition Elements, and Inner Transition Elements
Chemical Groups/Families
Alkali Metals
Shiny/Silver
Extremely reactive
Alkaline Earth Metals
Soft Metals
Reactive
Halogens
All states solid, liquid, and gas
Most reactive non-metal
Noble Gases
Gases
Inert (do not react)

Hydrogen

- Gas
- Reactive
- common in compounds (water)

Metalloids

- metallic and non-metallic properties

NOTE: The Chemical Reactivity of an Element is determined by its valence electrons all elements want a full valence shell.

Read p. 177 – 188
B16 – Exploring the Periodic Table
Learning Checkpoint p. 182 # 1 - 7
Learning Checkpoint p. 184 # 1 - 3
Learning Checkpoint p. 187 # 1 - 4
5.4 Check and Reflect # 1 – 14 (p. 188)
Chapter 5 Review p. 194 – 195 # 1 – 23
Unit B Review p. 199 – 201 #1 – 37
Review Supplemental (hand-out)